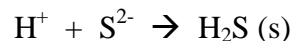
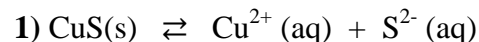


ANSWERS: Discuss factors affecting solubility

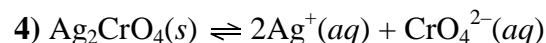


At a lower pH, $[\text{H}_3\text{O}^+]$ is increased so hydrogen sulfide is formed. This removes some of the S^{2-} product which causes the equilibrium to favour the forward reaction and the solid CuS moves into solution, so only the most insoluble sulfides in which the IP is still greater than the Ks (because the Ks is very small, re in CuS). ZnS's higher Ks means that it will not precipitate out.

But at a pH close to 7, hydrogen sulfide is not formed so the equilibrium strongly favours the back reaction and all metal sulfides precipitate.

2) When the pH is decreased, $[\text{H}_3\text{O}^+]$ will increase. The H_3O^+ will react with the OH^- and therefore remove them from the equilibrium. This will cause the reaction to replace some of the removed OH^- . As a result more Fe(OH)_3 will dissolve, so decreasing the pH will increase the solubility of Fe(OH)_3 .

3) Raising the pH will increase the concentration of OH^- ions. This will initially cause additional precipitate to form. Once the pH has been increased sufficiently (enough OH^- has been added) the formation of a complex ion with Zn^{2+} will occur, lowering OH^- ion concentration in solution. Thus the precipitate will redissolve as a complex ion and less precipitate will be at the bottom of the test tube.

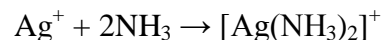


i) Less silver chromate dissolves.

Silver chromate is in equilibrium. More chromate ions shift this to the left.

ii) More silver chromate dissolves.

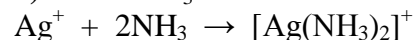
Ammonia complexes with silver ions.



This removes silver ions from the equilibrium mixture. The silver chromate equilibrium will shift to the right hand side or forward.

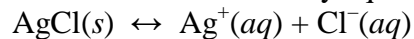
Note: for (c) (ii) accept NH_3 solution increases OH^- ion concentration so that Ag^+ reacts with this and precipitates out of solution (as Ag_2O). Hence, less Ag^+ in solution results in more Ag_2CrO_4 dissolving.

5) Dilute NH_3 will react with Ag^+ to produce the complex ion $[\text{Ag}(\text{NH}_3)_2]^+$.



This decreases $[\text{Ag}^+]$.

This moves the solubility equilibrium below to the right

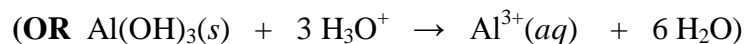
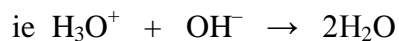


increasing the solubility of AgCl / causing more AgCl to dissolve,
therefore the precipitate disappears.



Acidic conditions

When the pH is less than 4, $[\text{H}_3\text{O}^+] > [\text{OH}^-]$ and the $[\text{H}_3\text{O}^+]$ ions react with OH^- ions from Al(OH)_3 .



The decrease in $[\text{OH}^-]$ from the solubility equilibrium causes the position of equilibrium to shift further to the right so that more Al(OH)_3 is dissolved.

Basic conditions

When the pH is greater than 10, $[\text{H}_3\text{O}^+] < [\text{OH}^-]$ and the OH^- ions react with Al^{3+} ions from $[\text{Al(OH)}_4]^-$



The decrease in $[\text{Al}^{3+}]$ from the solubility equilibrium causes the position of equilibrium to shift further to the right so that more Al(OH)_3 is dissolved.

7) The added Cl^- reduces the solubility of the NaCl. For the saturated solution:

$\text{NaCl}(s) \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq)$. The addition of Cl^- causes the equilibrium to favour the reactants and hence a precipitate will begin to form.